Interaction and Activation in Aqueous Solutions

Blandamer, M.J.; Engberts, J.B.F.N.

Published in:
Journal of the Chemical Society%2C Perkin Transactions 2

Publication date:
1995

INTERACTION AND ACTIVATION IN AQUEOUS SOLUTIONS

M. J. Blandamer and J. B. F. N. Engberts

Department of Chemistry, University of Leicester, LEICESTER-LE1 7RH, England
Department of Organic & Molecular Inorganic Chemistry, University of Groningen, Nijenborgh 4, 9747 AG GRONINGEN, The Netherlands

Received December 20, 1994

Abstract: The extent to which the properties of neutral solutes in dilute aqueous solution deviate from ideal can be related to pairwise Gibbs energy interaction parameters. These pairwise interaction parameters can be re-expressed in terms of group-group interaction parameters. Extension of these concepts to solute-initial state and solute-transition state interactions is reviewed in the context of understanding the effect of added solutes on rate constants for ester hydrolysis in aqueous solutions. Examples of application of the method of data analysis is reviewed and their significance discussed.

1. Introduction

In 1890 Menschutkin[1] reported that the rates of chemical reactions in solution are strongly influenced by the nature of the solvent. Since that time, the challenge of understanding the role of solvents in chemical reaction has attracted intense interest. A number of broad predictions were made by Ingold[2] on the basis of polarities of solvent, initial state and transition state. Reichardt[3] has shown the qualitative concepts of solvent polarity can be set down in quantitative form using E_T values based on spectrophotometric data. Nevertheless, within the multitude of potential solvent systems, aqueous systems stand out. Their biological and environmental importance justifies the attention given to aqueous systems. Moreover, the enormous amount of information concerning these systems should aid the development of satisfactory models for chemical reactions in solutions. In this context we are interested in two features of aqueous solutions. The term "hydration" describes solute-water interactions so that we imagine around a given solute-j a cosphere which encompasses all solvent molecules whose structure/organisation differs from that in the bulk pure solvent. Gurney[4] introduced the term cosphere to describe the water molecules perturbed by the solute-j. The second feature centres on deviations in the properties of neutral solute molecules in aqueous solutions from those defined as ideal properties (see below) which are understood in terms of cosphere-cosphere overlap.[5] In the analysis described below, we use a thermodynamic framework to describe these interactions[6] and then show how kinetic data can be analysed. Having linked kinetic data to a model using thermodynamics we indicate how the derived parameters can be interpreted.

2. Interactions in Aqueous Solutions

We envisage an aqueous solution prepared using 1 kg of water containing two solutes i and j, molalities m_i and m_j respectively. The chemical potential of solute-i, $\mu_i(aq)$ in this solution is given by equation (2-1).

$$\mu_i(aq) = \mu_i^o(aq) + RT\ln (m_i \gamma_i/m^o) \qquad (2\text{-}1)$$

Here $\mu_i^o(aq)$ is the standard (assuming ambient pressure is effectively the standard pressure) chemical potential of solute-i in an ideal aqueous solution, molality $m_i = m^o = 1$ mol kg^{-1}, at the same T and p. By definition, limit ($m_i \rightarrow 0$; $m_j \rightarrow 0$) γ_i equals unity at all T and p. A similar equation is written for solute-j in terms of $\mu_j^o(aq)$ and γ_j. Thus both γ_i and γ_j describe the extent to which solute-solute interactions account for the deviations in the properties of the real solution from ideal. In the present context we envisage for the moment that solute-i is a substrate (initial state). Then if solute-i is ideal (i.e. $\gamma_i = 1.0$), the chemical potential $\mu_i(aq)$ is unaffected by added solute-j. If $\gamma_i > 1.0$, the chemical potential of solute-i exceeds that in an ideal solution implying that, from a kinetic standpoint, solute-i is destabilised and thence more reactive. On the other hand, if $\gamma_i < 1.0$, solute-i is stabilised relative to the ideal solution implying that, from a kinetic standpoint, solute-i is less reactive (see below). [NB: γ_i cannot be negative.]

The chemical potential of the solvent $\mu_l(aq)$ is related to the chemical potential of pure water $\mu_l^*(l)$ at the same T and p using equation (2-2).

$$\mu_l(aq) = \mu_l^*(l) - \phi RTM_l(m_i + m_j) \qquad (2\text{-}2)$$

Here M_l is the molar mass of solvent water; ϕ is the practical osmotic coefficient which is unity in an ideal solution. Further, for an ideal solution, the chemical potential of the solvent is lower than that of the pure solvent. In fact, as we increase the total molality, $m_i + m_j$, so the stabilisation of the solvent water in an ideal solution increases; *cf.* osmotic pressures of solutions. In the context of a reaction involving direct participation of the solvent (see below), we anticipate that by increasing molalities of solutes the reactivity of the water decreases.

A key condition is offered by the Gibbs-Duhem equation which for the solution described above is written as follows (as fixed T and p).

$$(1/M_l)\, d\mu_l(aq) + m_j d\mu_j(aq) + m_i d\mu_i(aq) = 0 \qquad (2\text{-}3)$$

The equation is the key to much of the discussion which follows because it shows how a perturbation in the chemical potential of the solvent or one of the two solutes links through into the chemical potentials of all other substances in the system. If, for example, we use solute-i to model the initial state in a chemical reaction, then equation (2-3) indicates that $\mu_i(aq)$ is sensitive to changes in the chemical potential of solute-j. In other words, the Gibbs-Duhem equation concerns "communication" between all chemical substances in the system.

In summarising these key equations describing the chemical potentials of solvent and solutes, we anticipated their application to analysing kinetic data. A useful link in this context is achieved using the excess Gibbs enrgy G^E for the solution containing two solutes in 1 kg of water. Thus G^E expresses the difference between the Gibbs energies of the real and corresponding ideal solutions.

$$G^E(aq)/RT = (m_i + m_j)(1 - \phi) + m_j \ln \gamma_j + m_i \ln \gamma_i \qquad (2\text{-}4)$$

Using the Gibbs-Duhem equation we probe the change in G^E on adding a small amount of substance i at constant amounts of solvent and solute-j. Thus

$$\ln \gamma_i = (1/RT)[dG^E(aq)/dm_i] \qquad (2\text{-}5)$$

Hence if we can formulate an expression for G^E, we can obtain an equation for the activity coefficient of solute-i (see below).

3. Model Development

The underlying theme in the previous section concerns the extent to which the properties of a given solute in aqueous solution are sensitive to the nature and molality of another solute. Developing the theme, the scheme shown in Figure 1 emerges in which we consider the impact of added solute-j on the initial and transition states for a simple first order unimolecular reaction in solution involving solute-i.

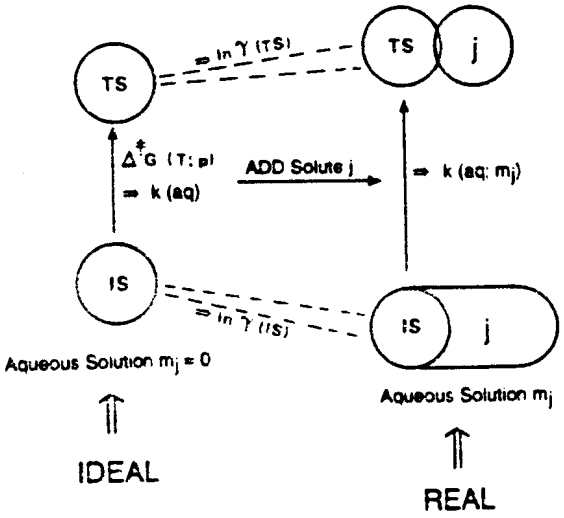
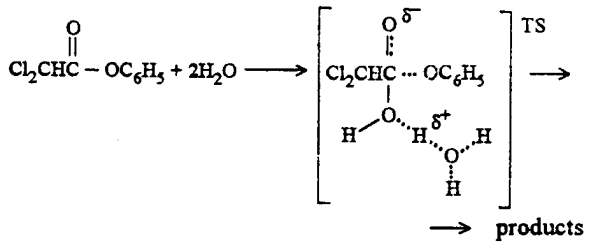


Figure 1. Diagrammatic representation of the effect of added solute-j on initial and transition states for a unimolecular-first order chemical reaction involving solute-i in aqueous solution; the pattern is one of the many possible - added solute-j stabilises the initial state and destabilises the transition state, hence producing a decrease in rate costant, i.e. an increase in $\Delta^{\neq}G^{o}$.

The arguments hinge on the extent to which overlap between (a) cospheres of initial state and solute-j and (b) cospheres of transition state and solute-j either stabilise or destabilise each state and hence change the standard activation Gibbs energy $\Delta^{\neq}G^{o}$. In this argument the emphasis is placed on the role of pairwise interaction between solute molecules in solution. The analysis developed below was prompted by observed patterns in salt effects on the rate constant for spontaneous hydrolysis of phenyl dichloroethanoate in aqueous solution[7] (at 298.2 K and ambient pressure); Scheme I. Rate constants were obtained (Figure 2) for spontaneous hydrolysis of the activated ester in a range of 1:1 salt solutions at constant salt concentration. Interestingly, the order of rate constants for a range of cations including hydrophilic alkali metal cations and hydrophobic tetraalkylammonium cations reversed on going from fluorides through chlorides to bromides. The pattern closely resembles that observed for the mean ionic coefficients for these salts at the same concentration.[8] In other words, the interactions between the solutes in

solution reflect the nature of their hydration characteristics and can be anaIysed from a quantitative standpoint.

Scheme1.



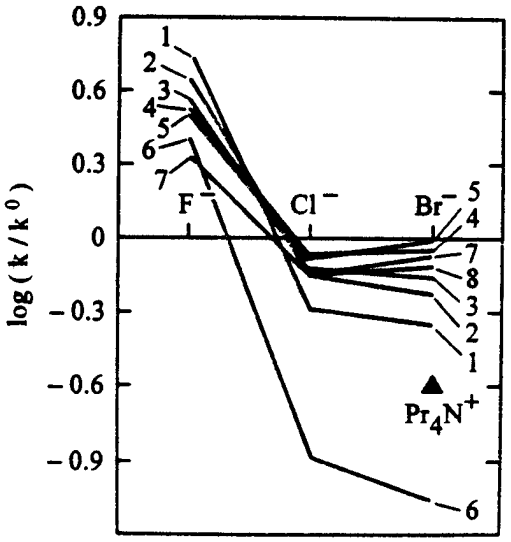


Figure 2. Dependence on salt (0.9 mol dm^{-3}; 298.2 K; ambient pressure) of the rate constant for hydrolysis in aqueous solution of phenyl dichloroethanoate: (1) Et_4N^+, (2) Me_4N^+, (3) K^+, (4) Cs^+, (5) Rb^+, (6) Bu_4N^+, (7) Na^+, (8) Li^+.

4. Interaction Parameter

Granted that the data summarised in Figure 2 highlight an interesting pattern, the next stage of the analysis requires a model linking the non-ideal properties of an aqueous solution with parameters describing solute-solute interaction parameters. For the solution described in section 2 containing solutes i and j, we express the excess Gibbs energy as a function of m_i and m_j using equation (4-1).

$$G^E(aq) = g_{jj}\,(m_j/m^o)^2 + 2g_{ij}m_im_j/(m^o)^2 + g_{ii}(m_i/m^o)^2 \qquad (4\text{-}1)$$

Eq (4-1) is satisfactory for dilute solutions where deviations from ideal can be described in terms of pairwise solute-solute interaction parameters, g_{jj}, g_{ij} and g_{ii}. Using the Gibbs-Duhem equation the activity coefficient γ_i is expressed in terms of g_{ii} and g_{ij}; eq (4-2).

$$\ln \gamma_i = (2/RT)(1/m^o)^2.[g_{ij}m_i + g_{ij}m_j] \qquad (4\text{-}2)$$

In the limit that there is only a trace amount of solute-i in the solution, we obtain an equation for the corresponding trace activity coefficient γ_i^T as a function of the molality of solute-j; eq (4-3).

$$\ln \gamma_i^T = (2/RT)g_{ij}(1/m^o)^2m_j \qquad (4\text{-}3)$$

In other words, $\ln \gamma_i^T$ is predicted to be a linear function of molality m_j, the proportionality constant being a function of g_{ij} which describes solute-i → solute-j interactions in aqueous solution. It should be noted that these pairwise interaction parameters do not necessarily signal that the two solute molecules are in contact. The argument developed above simply implies that the cospheres around pairs of solute molecules interact. In view of the fact that hydrogen bonding between water molecules has a strong cooperative element, we anticipate that the interaction between solute molecules "travels" through many intervening solvent molecules in aqueous solutions.

5. Solute-Solute Interactions and SWAG

The thermodynamic properties of aqueous solutions containing simple neutral molecules can be analysed to yield[9-11] pairwise Gibbs energy interaction parameters, *e. g.* g(ROH) ↔ g(ROH) for pairwise interaction between two monohydric alcohol molecules ROH in aqueous solution. An interesting development was suggested by Savage and Wood[12] who re-expressed these solute-solute interaction parameters in terms of pairwise group Gibbs energy interaction parameters; *e. g.* G(X ↔ Y) for groups X and Y – the SWAG model.[13] For example, in the context of monohydric alcohols G(CH_2 ↔ CH_2) is estimated at -23 J kg^{-1} for two interacting hydrophobic groups whereas G(OH ↔ CH_2) equals +29 J kg^{-1}, a repulsive interaction between two groups whose hydration characteristics are quite different.

6. Analysis of Kinetic Data

The procedures described above are brought together in an analysis of kinetic data based on the following arguments. We consider a substrate-i undergoing reaction in solution where the first order rate constant describes the following process.

$$\text{Solute-i} \leftrightarrow TS^{\neq} \rightarrow \text{products} \qquad (6\text{-}1)$$

Then based on classic transition state theory[14] the ratio of rate constants in real and ideal solutions, k/k^o

is related to the activity coefficients of transition state (≠) and initial state (i).

$$\ln (k/k^o) = -\ln \gamma_{\pm} + \ln \gamma_i \qquad (6\text{-}2)$$

In the aqueous solution the molality of added solute-j equals m_j but there are only trace amounts of substrate-i and transition state, ≠. If the chemical reaction involves n molecules of water on forming the transition state then eq (6-2) is rewritten to take account of the effect of added solute-j on the chemical potential of water. Thus

$$\ln (k/k^o) = -\ln \gamma_{\neq} + \ln \gamma_i - n\phi m_j M_l \qquad (6\text{-}3)$$

The two activity coefficients $\gamma_{\pm}$ and γ_i are re-expressed in terms of pairwise group interaction parameters, the latter being characteristic of groups u and v both substrate-i, transition state ≠ and added solute-j. Thus, when the solute molecules contain m- and k- different groups l and s,

$$\ln (k/k^o) = (2/RT)(1/m^o)^2 m_i \left\{ \left[\sum_{l=1}^{l=m} \sum_{s=1}^{s=k} n_l^{(i)} n_s^{(j)} G(l \leftrightarrow s) \right] - \left[\sum_{l=1}^{l=m} \sum_{s=1}^{s=k} n_l^{(\neq)} n_s^{(j)} G(l \leftrightarrow s) \right] \right\} - n\phi m_j M_l \qquad (6\text{-}4)$$

Eq (6-4) has formed the basis of several applications including the effects of added alcohols on the rate of hydrolysis of 1-acyl-1,2,4-triazoles in aqueous solutions.[15] An interesting example[16] concerns the effect of added urea and alkyl ureas on the rate constants for the neutral hydrolysis of p-methoxyphenyl dichloroethanoate;

Scheme 2.

p-CH_3O-C_6H_4-OCOCHCl_2

In this reaction two water molecules are incorporated from the solvent into the transition state. The ester is the solute-i and symbol U represents an added urea. Hence

$$\ln (k/k^o) = (2/RT)(1/m^o)^2 \{g(U \leftrightarrow i) - g(U \leftrightarrow \neq)\} m_u - 2\phi m_u M_l \qquad (6\text{-}5)$$

The key feature of this reaction is three polarised OH groups in the transition state. Further urea is equivalent to 1.5 CONH groups. Hence eq (6-5) is written in the following form.

$$\ln (k/k^o) = (2/RT)(1/m^o)^2\, 3(1.5) G(\text{CONH} \leftrightarrow \text{OH}^{\neq}) m_u - 2\phi m_u M_l \qquad (6\text{-}6)$$

Agreement between theory and experiment is quite satisfactory (see Figure 3).

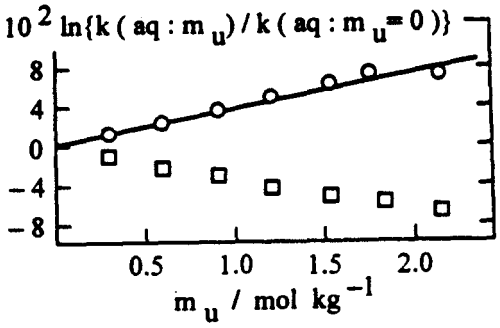


Figure 3. Dependence of ln (k/k^o) for hydrolysis of p-methoxyphenyl dichloroethanoate in aqueous solution as a function of molality of added urea; experimental points (O) and calculated dependence (full line) using G(CONH ↔ $OH^{\neq}$) = -20 J kg^{-1}. The points (❑) are calculated with G(CONH ↔ $OH^{\neq}$) set to zero, the trend being a consequence of added urea on chemical potential of water in urea (aq).

The dependence of rate constant for the same reaction on molality of added dimethyl urea can be accounted for by incorporating into the analysis the pairwise interaction parameter G(CH_2 ↔ $OH^{\neq}$) describing the interaction between methylene groups and the polarised OH-groups in the transition state. Significantly, having estimated the interaction parameters G(CONH ↔ $OH^{\neq}$) and G(CH_2 ↔ $OH^{\neq}$) no additional information is required to account for the dependence of rate constant for the same reaction on the molality of added tetramethyl urea (Figure 4). The agreement between experimental and predicted dependences offers considerable support to the analytical procedures discussed in the previous sections.

The treatment described here has been used in the analysis of kinetic data for reactions in a range of other systems and reactions.[17] The success has been encouraging. Two interesting and recent examples[21,22] concern the hydrolysis of 1-benzoyl-3-phenyl-1,2,4-triazole (Scheme III) and p-methoxyphenyl dichloroethanoate in aqueous solutions containing carboxamides. The dependence of rate constant on molality of carboxamide was expressed using eq (6-7) where G(C) is the Gibbs energy interaction parameter.

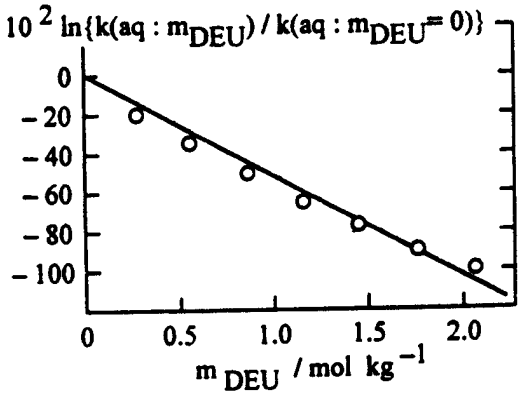


Figure 4. Dependence of ln (k/k^o) for hydrolysis of p-methoxyphenyl dichloroethanoate in aqueous solution as a function of the molality of added diethylurea; experimental points (O) and calculated dependence with pairwise interaction parameters calculated from the corresponding dependence on molality of urea and dimethy1urea; $G(CONH \leftrightarrow OH^{\neq}) = -20$ J kg^{-1} and $G(CH_2 \leftrightarrow OH^{\neq}) = 45$ J kg^{-1}.

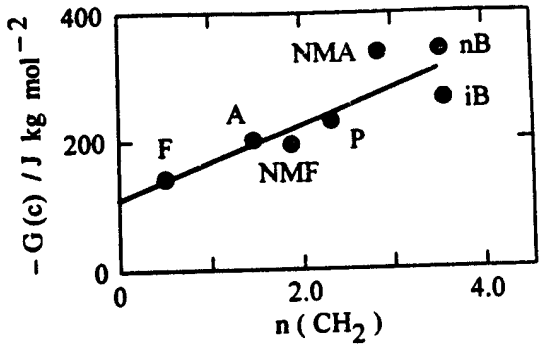
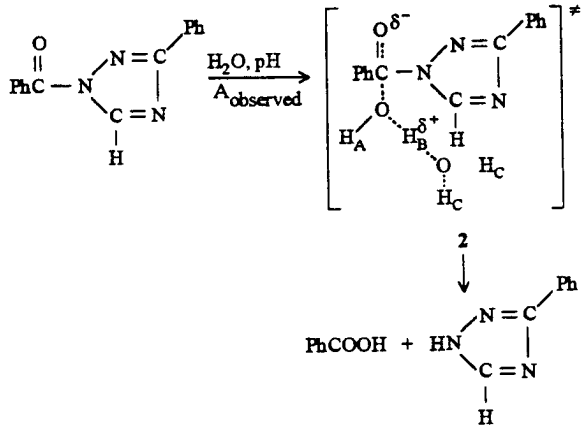


Figure 5. Dependence of Gibbs energy interaction parameter G(C) on the number of methylene groups in formamide (F), acetamide (A), N-methyl formamide (NMF), propionamide (P), N-methyl acetamide (NMA), n-butyramide (nB) and iso-butyramide (iB) calculated trom the dependence of rate constant for the hydrolysis of 1-benzoyl-3-phenyl-1,2,4-triazole on molality of added carboxamide.

Scheme 3.



The model outlined above is supported by the observation that G(C) is a linear function of the number of methylene groups in the added solute-j;

$$\ln (k/k^o) = (2/RT)G(C)m_j - n\phi m_j M_l \qquad (6\text{-}7)$$

7. Conclusions

One task of thermodynamics is to build a bridge between models for systems and measured properties of systems. Here the task has been to link kinetic data with models describing *interactions* between solutes in aqueous solutions. The results described here show that an analysis based on pairwise solute-solute and pairwise group-group interaction parameters has obvious merit despite the well-established complexity of aqueous solutions.

References

1. N. Menschutkin, *Z. Phys. Chem.*, **5**, 589 (1890).
2. C. K. Ingold, *Structure and Mechanism in Organic Chemistry*, Bell, London, 1953.
3. C. Reichardt, S. Asharin-Fard, A. Blum, M. Eschner, A.-M. Mehranpour, P. Milarr, T. Niem, G. Schäfer, M. Wilk, *Pure Appl. Chem.*, **65**, 2593 (1993).
4. R. W. Gurney, *Ionic Processes in Solution*, McGraw-Hill, New York, 1953.
5. H. L. Friedman, C. V. Krishan, *Water - A Comprehensive Treatise* (ed. F. Franks), Plenum Press, New York, Vol.III, Chapter 1., 1973.
6. M. J. Blandamer, J. Burgess, J. B. F. N. Engberts, W. Blokzijl, *Annual Reports* (*Royal Society of Chemistry - Sect.C*), 45 (1988).
7. J. B. F. N. Engberts, M. J. Blandamer, J. Burgess, B. Clark, A. W. Hakin, *J. Chem. Soc., Faraday Trans.1*, **83**, 865 (1987).
8. J. E. Desnoyers, M. Arel, G. Perron, C. Jolicoeur, *J. Phys. Chem.*, **73**, 3346 (1969).
9. S. K. Suri, R. H. Wood, *J. Soln. Chem.*, **15**, 705 (1986) (and references therein).
10. T. Lilley, *Pure Appl. Chem.*, **65**, 2551 (1993) (and references therein).
11. G. Della Gatta, G. Barone, V. Elia, *J. Soln. Chem.*,**15**, 157 (1986) (and references therein).

12. J. J. Savage, R.H. Wood, *J. Soln. Chem.*, **5**, 33 (1976).
13. SWAG = Savage-Wood Additivity of Group Interactions.
14. S. Glasstone, K. J. Laidler, H. Eyring, *The Theory of Rate Processes*, McGraw-Hill, New York, 1941.
15. W. Blokzijl, J. Jager, J. B. F. N. Engberts, M. J. Blandamer, *J. Amer. Chem. Soc.*, **108**, 6411 (1986).
16. W. Blokzijl, J. B. F. N. Engberts, J. Jager, M. J. Blandamer, *J. Phys. Chem.*, **91**, 6022 (1987).
17. H. Benak, J. B. F. N. Engberts, M. J. Blandamer, *J. Chem. Soc., Perkin Trans. 2*, 2035 (1992).
18. W. Blokzijl, J. B. F. N. Engberts, M. J. Blandamer, *J. Chem. Soc., Perkin Trans. 2*, 455 (1994).
19. S. A. Galema, M. J. Blandamer, J. B. F. N. Engberts, *J. Amer. Chem. Soc.*, **112**, 9665 (1990).
20. W. Blokzijl, J. B. F. N. Engberts, M. J. Blandamer, *J. Amer. Chem. Soc.*, **113**, 4241 (1991).
21. R. P. V. Kerstholt, J. B. F. N. Engberts, M. J. Blandamer, *J. Chem. Soc., Perkin Trans. 2*, 49 (1993).
22. J. B. F. N. Engberts, R. P. V. Kerstholt, M. J. Blandamer, *J. Chem. Soc., Chem. Commun.*, 1230 (1991).